

A2.1

What substances can act as sunscreens?

In this activity, you will use information about the absorption of visible and ultraviolet radiation by some substances to decide how they will affect radiation from the Sun.

Look at Figure 1, which shows the portion of the Sun's radiation that reaches us on Earth, and its effect on skin.

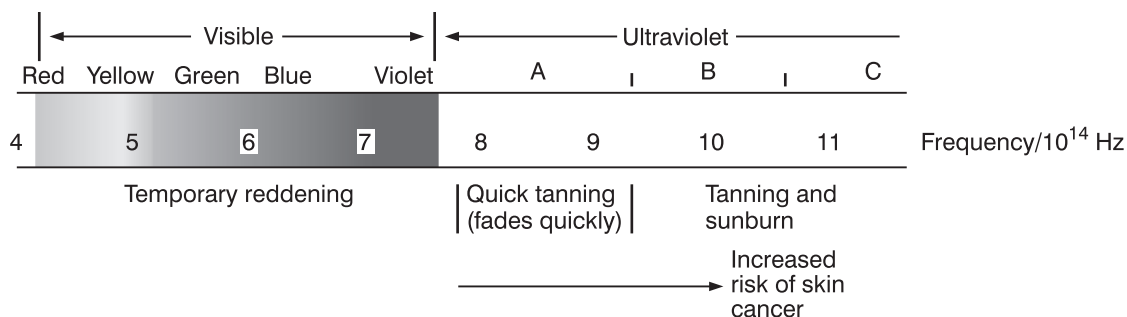


Figure 1 The portion of the Sun's radiation that reaches the Earth and its effect on the skin (from **The Atmosphere, Storyline A2**)

Each of the substances listed in Table 1 absorbs light in the visible–ultraviolet region. The **absorption spectrum** of each substance shows the intensity of radiation absorbed at each frequency, and may consist of either a number of sharp peaks or a broad band of absorption called a continuum. The figures in Table 1 show the lowest frequency (or frequency range) at which each substance absorbs strongly.

Substance	Absorption frequency range/ 10^{14} Hz
water	from 16.2
glass	from 9.4
glyceryl trioleate (a major constituent of olive oil)	6.8–10.7
4-aminobenzoic acid	8.8–12.2

Table 1 Ultraviolet absorption of some substances

Where possible, mark on Figure 1 the start of the absorption range of the substances in the table above. (Some of them may lie outside the frequency range shown in the diagram.)

What are your conclusions concerning:

- whether swimming can protect you from sunburn?
- the chance of getting tanned sitting in a south-facing window on a sunny day?
- the use of olive oil as a sunscreen?
- the use of 4-aminobenzoic acid as a sunscreen?

This activity will help you develop the important skill of designing and carrying out a practical investigation.

Requirements

- a light source which provides radiation of about the same frequency as the ultraviolet light that causes sunburn and skin cancer
- a means of supporting the sunscreen so that light can be directed at it and possibly pass through it
- a way of detecting the ultraviolet light
- a variety of sunscreen products with different screening factors

CARE If you are using any kind of ultraviolet lamp, remember that **ultraviolet radiation is hazardous**. Follow the recommended precautions concerning eye protection.



What you do

The aim of this activity is to find out about the effectiveness of different sunscreens.

The design of the experiment is left to you. You will need to think about the items in the requirement list, and decide how you are going to test the sunscreens so that you can make fair comparisons. You may decide on modifications after your first try which lead to improvements in your method.

Prepare a **Risk Assessment** for your planned activity and a list of sources you have used in developing your plan and your Risk Assessment.

Do not start your investigation until your plan has been checked by your teacher.

The following points may help you.

- 1 The best source of ultraviolet radiation of the right frequencies is of course the Sun itself. A domestic sun-tan lamp is also good, but does not emit frequencies above about 9.3×10^{14} Hz. A mid-range ultraviolet lamp from the school science department will emit frequencies in the 9.3×10^{14} Hz – 10.7×10^{14} Hz region, where the skin is very sensitive.
- 2 Most ultraviolet radiation will not pass through glass. It passes quite well through perspex, and very well through clingfilm.
- 3 For detecting the ultraviolet radiation, there are a number of possibilities:
 - a Ultraviolet sensitive paper, which turns blue in the presence of natural ultraviolet radiation. (Incidentally, this paper is used to make security passes to check whether the wearer has left the building. You can use it to test the effectiveness of glass as a sunscreen by sticking pieces inside and outside a window.)
 - b A white cloth that has been washed in detergent. (Most machine detergents contain fluorescent compounds. You can see the effect under some disco lights.) The cloth will give a bright fluorescence with ultraviolet light.
 - c Tonic water fluoresces in ultraviolet light.
 - d Some photocopying paper fluoresces.

A2.3

What is the effect of the atmosphere on the Sun's radiation?

In this activity, you will use information about the ultraviolet absorption of some atmospheric gases to investigate their effect on radiation from the Sun.

Figure 1 shows the spectrum of the Sun's radiation in space, that is before it has passed through the Earth's atmosphere. It shows the energy emitted at each wavelength (and frequency).

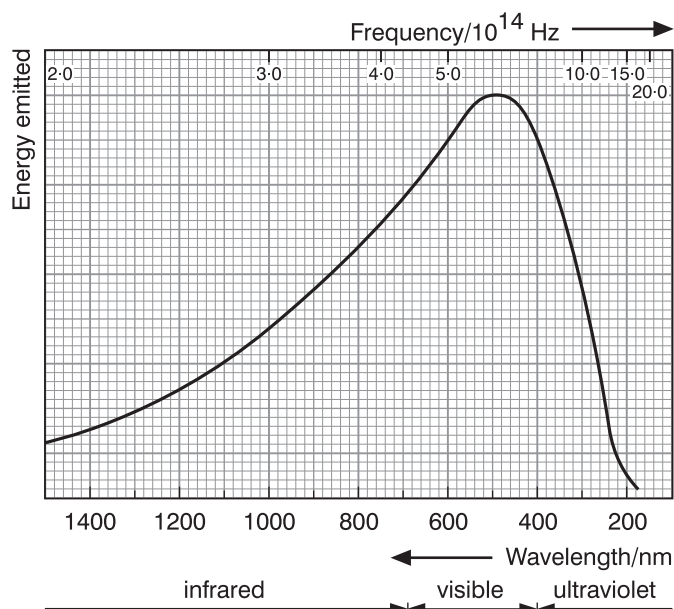


Figure 1 The spectrum of solar radiation in space

Table 1 shows the frequency and wavelength ranges over which some common atmospheric gases absorb ultraviolet radiation. Where only one figure is given, it represents the low frequency (high wavelength) limit at which absorption begins.

You will find it easier to use the wavelength ranges rather than the frequency ranges. This is because the horizontal axis of the spectrum is plotted using a linear wavelength scale, whereas the frequency is non-linear.

Gas	Absorption frequency range/ 10^{14} Hz	Absorption wavelength range/nm
oxygen	12.4–17.1 (very weak) and above 17.1 (strong)	242–175 (very weak) and below 175 (strong)
nitrogen	above 50.0	below 60
water	16.2–20.6	185–145
carbon dioxide	17.6–23.9 (weak) and above 25.6	170–125 (weak) and below 117
ozone	10.1–14.0	296–214
methane	above 20.8	below 144
sulphur dioxide	9.1–12.5 (weak) and above 13.8	330–240 (weak) and below 217

Table 1 Ultraviolet absorption of some atmospheric gases

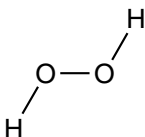
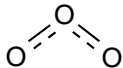
- On Figure 1 shade the area of the solar spectrum which corresponds to ultraviolet radiation.
- List the gases from Table 1 that absorb in this region, with the relevant absorption ranges. Mark and label these regions on Figure 1 using a double-headed arrow \leftrightarrow .
 - Comment on the effect of each of the gases in the atmosphere on solar radiation. You should consider both the absorption characteristics and the concentrations of the gases in the atmosphere.
 - How would you expect the Sun's spectrum as seen on the surface of the Earth to differ from that seen in space?

In this activity, you can find out more about the bonding in ozone and the stability of ozone relative to oxygen. You can remind yourself about the use of bond enthalpies to calculate enthalpy changes.

Ozone is a polymorph (or allotrope) of oxygen. Its formula is O_3 and it is called trioxygen. It is a pungent-smelling, pale blue gas which condenses at low temperatures to give a dark blue liquid. It is much less abundant in the atmosphere than the other polymorph, dioxygen, O_2 .

Bonding in ozone

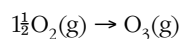
The table shows three substances which contain oxygen–oxygen bonds.

Substance	Shape of molecule	Length of O–O bond/nm	Bond enthalpy of O–O bond /kJ mol ⁻¹
dioxygen, O_2 (double bond)	$O=O$		
hydrogen peroxide, H_2O_2 (single bond)			
ozone, O_3			

- Use the **Data Sheets** to fill in the bond length and bond enthalpy of the oxygen–oxygen bond in each compound.
- What do these figures tell you about the nature of the bonding in ozone?

Energetics

The overall equation for the formation of ozone from dioxygen is:



- Work out the standard enthalpy change of formation of ozone by the following steps.
 - Draw a Hess cycle to show the formation of ozone from dioxygen via three atoms of oxygen in the gaseous state.
 - Use the bond enthalpies in your table to work out a value for the standard enthalpy change of formation of ozone. (Remember that the bond enthalpy is the energy needed to break one mole of bonds, see **Chemical Ideas 4.2.**)
- Which is the more energetically stable allotrope of oxygen, O_2 or O_3 ? Explain your answer.
- Ozone is described as an *endothermic substance*. What does this mean?
- Would you expect ozone to be more or less reactive than dioxygen? Explain your answer.

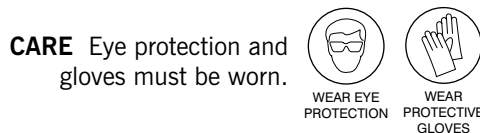
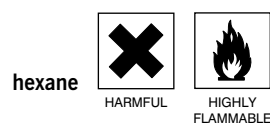
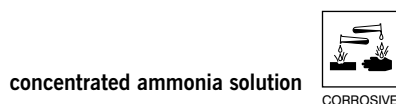
The photodissociation of bromine (Optional teacher demonstration)

In this activity you will investigate the effect of light on the reaction between bromine and hexane. You will learn how to handle bromine safely.

Requirements

- hexane (6 cm^3)
- bromine (3 drops)
- aluminium foil
- test-tubes and rack
- Universal Indicator paper
- glass rod
- concentrated ammonia solution
- light source (sunlight is best, but you could use a strong lamp such as a fluorescent light tube – a microscope lamp or an overhead projector lamp will work reasonably well)

CARE Bromine liquid should always be used in a fume cupboard. It produces an irritant vapour which is very toxic. The liquid causes severe burns. A beaker of sodium thiosulphate solution (approximately 1 mol dm^{-3}) should be available to treat spillages.



What you do

- 1 Put 2 cm^3 of hexane in each of three test-tubes. (**CARE** Hexane is highly flammable and harmful. Avoid skin contact and do not breathe the vapour.)
- 2 Add one drop of bromine (not bromine water) to each tube. (**CARE** Bromine is corrosive and toxic. Wear eye protection and gloves. Work in a fume cupboard when adding the bromine.)
- 3 Wrap aluminium foil around two of the tubes as shown in Figure 1. Stand the three tubes side by side in a test-tube rack.
- 4 Leave the rack in bright sunlight or next to a bright light source for 5–10 minutes, then examine the appearance of each tube. Cautiously blow across the top of each tube. Test any gases given off with moist indicator paper, and by holding a drop of ammonia solution on a glass rod at the mouth of the test-tube.

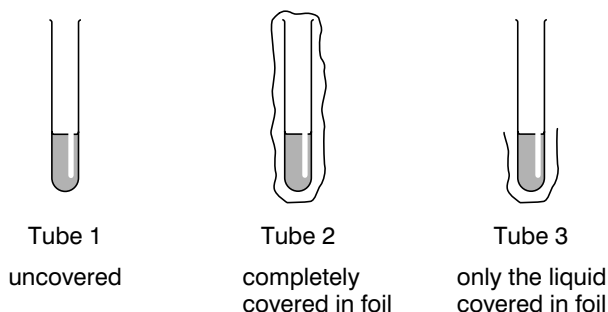


Figure 1

QUESTIONS

- a What evidence is there that bromine undergoes photodissociation?
- b Look up the values for Br–Br and C–H bond enthalpies. Remember that bond enthalpies are given per mole of bonds. Calculate the enthalpy change when:
 - i one Br–Br bond is broken
 - ii one C–H bond is broken.

Use $E = h\nu$ to calculate the frequencies of photons of radiation corresponding to each of these energies. Which bond is most likely to be broken by absorption of sunlight?
- c What has the bromine reacted with?
- d What is the gaseous product that is formed?
- e Write an equation, with state symbols, for the reaction of this gas with ammonia.









Investigating the reaction between bromine and hexane

In this activity you will carry out a quantitative investigation of the reaction between bromine radicals and hexane, and determine the amount of one of the products formed. This will give you an opportunity to improve your skills at carrying out an accurate titration, and to use the results of a titration to calculate the information needed.

Requirements

- reagent bottle (about 250 cm³ capacity) with good-fitting stopper
- hexane (35 cm³)
- protective gloves
- distilled water
- bromine (1 g)
- marked pipette for adding bromine
- sodium hydroxide solution, 0.200 mol dm⁻³ (100 cm³)
- Phenolphthalein indicator solution
- burette and stand
- wash bottle with distilled water
- access to a balance
- light source (sunlight is best, but you could use a strong lamp such as a fluorescent light tube – a microscope lamp or an overhead projector lamp will work reasonably well)

CARE Bromine liquid should always be used in a fume cupboard. It produces an irritant vapour which is very toxic. The liquid causes severe burns. A beaker of sodium thiosulphate solution (approximately 1 mol dm⁻³) should be available to treat spillages.

bromine		
	TOXIC	CORROSIVE
hexane		
	HARMFUL	HIGHLY FLAMMABLE
dilute sodium hydroxide solution		
	IRRITANT	
Phenolphthalein indicator solution		
	HIGHLY FLAMMABLE	
CARE Eye protection and gloves must be worn.		
	WEAR EYE PROTECTION	WEAR PROTECTIVE GLOVES

Outline of the experiment

Bromine is added to a mixture of hexane and water, and the mixture is placed in a bright light and shaken from time to time. Hydrogen bromide (HBr) is formed, and passes into the water layer.

When all the bromine has reacted and the red colour has gone, the mixture is titrated with sodium hydroxide. HBr(aq) reacts with the sodium hydroxide, and from the titration result the amount of HBr formed per mole of bromine can be calculated.

What you do

- 1 Put 35 cm³ of hexane and 15 cm³ of distilled water in the bottle and stopper it firmly. (**CARE** Hexane is highly flammable and harmful. Avoid skin contact and do not breathe the vapour.)
- 2 Make sure the bottle is dry on the outside, then find the mass of the bottle, stopper and contents.
- 3 Add approximately 1 g (about 0.3 cm³) of bromine to the bottle. (**CARE** Bromine is corrosive and toxic. Wear eye protection and gloves. Work in a fume cupboard when adding the bromine. Do not remove the stopper from the bottle after the bromine has been added.)
- 4 Find the mass of the bottle again, in order to find the exact mass of bromine added.
- 5 Shake the bottle. Place it near a light source or on a window sill. Shake the bottle every 2 minutes until the colour has disappeared.

- 6 Fill a clean burette with $0.200 \text{ mol dm}^{-3}$ sodium hydroxide solution. (**CARE** Sodium hydroxide of this concentration is an irritant.)
- 7 Give the stoppered bottle a final shake, remove the stopper and rinse it with distilled water into the bottle so no liquid is lost.
- 8 Add 10 drops of Phenolphthalein to the bottle. Titrate the contents of the bottle with $0.200 \text{ mol dm}^{-3}$ sodium hydroxide solution to a faint pink end-point. The end-point is not very stable and the colour tends to fade. Titrate to the first *overall* pink colour. (**CARE** The product of the reaction is flammable. Its vapours are harmful and may irritate the eyes and nose. Avoid breathing the vapours.)
- 9 When you have finished, the reaction mixture must be poured into a residue bottle *in a fume cupboard*.

Using the results

Record your results in a suitable table, then answer the following questions.

- a What mass of bromine did you use?
- b What amount in moles of bromine, Br_2 , did you use?
- c What volume of sodium hydroxide solution was needed in the titration?
- d What amount in moles of sodium hydroxide, NaOH , was used in the titration?
- e Write an equation for the reaction between sodium hydroxide and hydrogen bromide.
- f What amount in moles of hydrogen bromide was produced in the reaction between bromine and hexane?
- g What amount in moles of hydrogen bromide would be produced from the reaction of 1 mole of bromine, Br_2 ?
- h Write a balanced equation for the reaction between hexane and bromine to produce bromohexane, $\text{C}_6\text{H}_{11}\text{Br}$.
- i What general type of reaction is this?
- j Why was water added to the reaction mixture, even though it is not a reactant?
- k Bromohexane is formed in this reaction. In which of the two layers would it be found?

Evaluating your results and procedures

- l For which of your measurements is the level of precision you could achieve important? Work out the percentage error for each of these measurements.
- m Compare the relative importance of each of these errors.
- n Identify the stages in your *procedure* which could have led to errors.
- o How do you think the sources of error or uncertainty that you have identified in l and n affect the accuracy and reliability of your overall result?

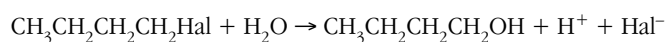
How do halogenoalkanes differ in reactivity?

This part of the activity will allow you to practise and develop your skills in experimental design. You are asked to plan an experiment to compare the reactivity of some halogenoalkanes.

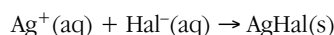
Some background information

Halogenoalkanes have the general formula RHal, where R is an alkyl group and Hal is a halogen atom. In this experiment you will investigate how the reactivity of halogenoalkanes depends on the nature of the halogen atom, Hal.

You will compare the rates of hydrolysis (reaction with water) of 1-chlorobutane, 1-bromobutane and 1-iodobutane.



The rate of the reaction is followed by carrying it out in the presence of silver ions, $\text{Ag}^+(\text{aq})$. Silver ions react with halide ions to form a precipitate of silver halide:



Halogenoalkanes are covalently bonded, so they give no precipitate with silver ions. But as the reaction proceeds and halide ions are produced, a white or yellow precipitate of silver halide gradually appears.

Halogenoalkanes are insoluble in water, so the reaction is carried out in the presence of ethanol, which acts as a mutual solvent for the halogenoalkane, the water and the silver ions.

Planning

Write a plan for how you will compare the rate of hydrolysis of a few drops of the three halogenoalkanes by warm water. Your plan should include the following:

- the apparatus to be used
- the quantities of materials to be used
- details of what you expect to do in the form of a set of instructions that could be followed by another student without any further guidance
- a **Risk Assessment** which identifies the hazards associated with each substance you will use and each operation you will carry out, and what steps you will take to minimise the hazards
- details of the sources you have used in devising your plan and your Risk Assessment.

Do not carry out your plan until it has been checked by your teacher.

A4.1b

How do halogenoalkanes differ in reactivity?

In this part of the activity, you will carry out the experiment to compare the reactivity of some halogenoalkanes. From your results, you should be able to see why Midgley was led to look at CFCs for a refrigerant gas.









Requirements

- test-tubes
 - ethanol (3 cm³)
 - silver nitrate solution, 0.01 mol dm⁻³ (3 cm³)
 - 0–110 °C thermometer
 - 250 cm³ beaker
 - source of hot water (50 °C)
 - 1-chlorobutane
 - 1-bromobutane
 - 1-iodobutane
- } a few drops of each
- stopclock

CARE Eye protection must be worn.



WEAR EYE PROTECTION

1-bromobutane	 HARMFUL	 HIGHLY FLAMMABLE
1-chlorobutane	 HARMFUL	 HIGHLY FLAMMABLE
ethanol	 HARMFUL	 HIGHLY FLAMMABLE
1-iodobutane	 HARMFUL	 FLAMMABLE

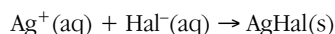
Some background information

Halogenoalkanes have the general formula RHal, where R is an alkyl group and Hal is a halogen atom. In this experiment you will investigate how the reactivity of halogenoalkanes depends on the nature of the halogen atom, Hal.

You will compare the rates of hydrolysis (reaction with water) of 1-chlorobutane, 1-bromobutane and 1-iodobutane.



The rate of the reaction is followed by carrying it out in the presence of silver ions, Ag⁺(aq). Silver ions react with halide ions to form a precipitate of silver halide:



Halogenoalkanes are covalently bonded, so they give no precipitate with silver ions. But as the reaction proceeds and halide ions are produced, a white or yellow precipitate of silver halide gradually appears.

Halogenoalkanes are insoluble in water, so the reaction is carried out in the presence of ethanol, which acts as a mutual solvent for the halogenoalkane, the water and the silver ions.

What you do

- 1 Heat some water to about 50 °C in a beaker. Extinguish the flame. (Alternatively, the water can be heated safely using an electric kettle.)
- 2 Set up three test-tubes labelled **1**, **2**, **3**. Place 1 cm³ of ethanol into each. (**CARE** Ethanol is highly flammable. Keep bottles stoppered when not in use and well away from naked flames. Avoid skin contact and do not breathe the vapour.)
- 3 Add 2 drops of chlorobutane to tube **1**, 2 drops of bromobutane to tube **2**, and 2 drops of iodobutane to tube **3**. (**CARE** Halogenoalkanes are flammable. Their vapours are irritating and harmful. Avoid inhaling them.)
- 4 Stand the tubes in the beaker of water at 50 °C. Put three test-tubes, each containing 1 cm³ of 0.01 mol dm⁻³ silver nitrate solution, in the same beaker. Leave the tubes for about 10 minutes so that they reach the temperature of the bath.
- 5 Now add 1 cm³ of the warm silver nitrate solution to each of tubes **1**, **2** and **3**, working quickly and noting the time. Shake each tube to mix the contents.
- 6 Observe the tubes over the course of the next 5 minutes or so and note down the time when any precipitate appears.

RESULTS AND QUESTIONS

If a precipitate appears it means that hydrolysis has taken place and halide ions have been released from the halogenoalkane.

- a Which halogenoalkane undergoes the fastest hydrolysis? Which is slowest?
- b Suggest a reason for the different rates of hydrolysis. The chart showing bond enthalpies in the **Data Sheets** may help you.
- c What result would you predict for 1-fluorobutane? Explain your answer.
- d Why was the experiment done using halogenobutanes rather than halogenomethanes?
- e What bearing do your results have on the choice of the best halogenoalkanes to use as refrigerant fluids?

In this activity, you will prepare a sample of 2-chloro-2-methylpropane. This will involve learning new techniques both to carry out the reaction, and to separate and purify the liquid product. You will need to work carefully to obtain the maximum yield possible.

Requirements

- 10 cm³ measuring cylinder
- 50 cm³ measuring cylinder
- 2-methylpropan-2-ol (6.5 cm³)
- concentrated hydrochloric acid (20 cm³)
- access to a balance
- 50 cm³ pear-shaped flask
- 50 cm³ separating funnel and stopper
- distillation head
- clamps and stand
- anti-bumping granules
- condenser
- small Bunsen burner or electric heating mantle
- sodium hydrogencarbonate solution, 5% (10 cm³)
- anhydrous sodium sulphate
- 100 cm³ conical flask (2)
- 0–110 °C thermometer and holder
- specimen tube
- 50 cm³ beaker

concentrated hydrochloric acid



CORROSIVE

2-methylpropan-2-ol



HIGHLY
FLAMMABLE



HARMFUL

CARE Eye protection
must be worn.



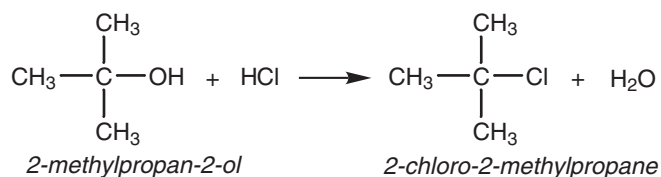
WEAR EYE
PROTECTION

CARE The vapours produced in this activity can be harmful. Work in a well-ventilated laboratory and avoid inhaling the fumes.

Background

One way to make a halogenoalkane is to start with an alcohol and replace the –OH group by a halogen atom.

For example, you can make 2-chloro-2-methylpropane by allowing 2-methylpropan-2-ol to react with concentrated hydrochloric acid at room temperature. The overall reaction can be represented by the equation:



The mechanism of this reaction is explained in **Chemical Ideas 14.2**. (A similar type of reaction with HBr rather than HCl is described in **Chemical Ideas 13.1**.)

The preparation of an organic compound usually takes place in four stages:

- carrying out the reaction
 - separating the required product from the reaction mixture
 - purifying the product
 - testing the product to check that it is a pure sample of the required compound.
- a Write the equation for the reaction given above using skeletal rather than structural formulae.

What you do _____

Carrying out the reaction

Before you carry out the reaction you need to learn how to use the separating funnel correctly. Your teacher will show you.

- 1 Into a 10 cm³ measuring cylinder, pour about 6.5 cm³ of 2-methylpropan-2-ol. (**CARE** Highly flammable and harmful. Keep bottle stoppered when not in use and well away from naked flames. Avoid skin contact and do not breathe the vapour.) This will be approximately 5 g of the 2-methylpropan-2-ol. Weigh the measuring cylinder and contents and then pour the 2-methylpropan-2-ol into a 50 cm³ separating funnel. Weigh the empty measuring cylinder and record the exact mass of 2-methylpropan-2-ol you have added to the funnel.
- 2 Into a 50 cm³ measuring cylinder, pour about 20 cm³ of concentrated hydrochloric acid. (**CARE** Extremely corrosive. Avoid skin contact.) Gradually add the acid to the 2-methylpropan-2-ol in the funnel over a period of about 2 minutes.
- 3 Put the stopper in the funnel and shake the mixture from time to time over the next 20 minutes. After each shaking *remove the stopper briefly to release the pressure*.

Separating the product from the reaction mixture

The mixture now contains the product 2-chloro-2-methylpropane (**CARE** Highly flammable and harmful) and the following substances:

- 2-methylpropan-2-ol
- hydrogen chloride (hydrochloric acid)
- water.

- 4 Allow the mixture to stand until two layers have separated (Figure 1). Remove the stopper and run the lower layer into a clean 100 cm³ conical flask. Use the table of densities (Table 1) to decide which layer contains the 2-chloro-2-methylpropane.

Liquid	Density/g cm ⁻³
2-chloro-2-methylpropane	0.84
2-methylpropan-2-ol	0.78
water	1.0
concentrated hydrochloric acid	1.2

Table 1 Densities of some of the liquids present

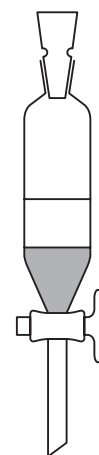


Figure 1 Separating funnel

Most of the 2-methylpropan-2-ol impurity dissolves in the concentrated hydrochloric acid layer and so is removed when you run off this layer. The remaining impurities are removed in the operations which follow.

- b Suggest why the alcohol dissolves in concentrated hydrochloric acid.
 - c What impurities are likely still to contaminate the 2-chloro-2-methylpropane layer at this stage?
- 5 *Slowly* add 10 cm³ of 5% sodium hydrogencarbonate solution to the 2-chloro-2-methylpropane in the separating funnel. Stopper the funnel and shake the contents gently and then more vigorously. *There will be a marked build up of gas, so be very careful when you release the pressure*. When the two layers have separated, run off and discard the lower aqueous layer. Repeat the washing with sodium hydrogencarbonate solution until no more gas is given off.
 - 6 Add 10 cm³ of distilled water and shake. When the two layers have separated, run off and discard the lower aqueous layer and then run the 2-chloro-2-methylpropane into a clean conical flask.
 - 7 To the liquid in the flask, add a small amount of anhydrous sodium sulphate, which acts as a **drying agent** and removes the last traces of water. Add the drying agent in small quantities, swirling after each addition, until the liquid is totally clear.

- d Which impurity is removed by shaking the product with a solution of sodium hydrogencarbonate?
- e Why is there a marked build up of pressure during the shaking?
- f Suggest a reason why sodium hydrogencarbonate solution is used rather than sodium hydroxide solution.
- g How did the appearance of your product change when it was swirled with anhydrous sodium sulphate? How do you account for this?
- h Name another drying agent which could be used in place of the anhydrous sodium sulphate.

Purifying the product and testing its identity and purity

Most of the impurities should now have been removed, except for a small amount of 2-methylpropan-2-ol which will be dissolved in the 2-chloro-2-methylpropane. This can be separated by carrying out a **distillation**.

- 8 Set up clean and dry apparatus for a simple distillation, complete with thermometer (see Figure 2). The thermometer bulb should be opposite the side-arm so that it measures the temperature of the liquid that distils over.

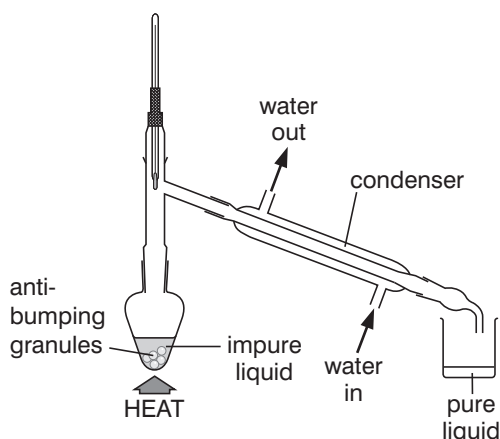


Figure 2 Apparatus for a simple distillation

- 9 Transfer the dried 2-chloro-2-methylpropane into the distillation flask and add a few anti-bumping granules.
- 10 Weigh a clean, dry specimen tube. *Gently* heat the liquid in the distillation flask using the flame from a small hand-held Bunsen burner. At first use a small beaker as the receiver. Start collecting the 2-chloro-2-methylpropane in the weighed specimen tube when the temperature reaches 48 °C. The boiling point of 2-chloro-2-methylpropane is 51 °C. Stop the distillation when the thermometer rises above 53 °C. Stopper the specimen tube and record the mass of 2-chloro-2-methylpropane collected.
- i How does the final stage allow you to check the identity and purity of your product?
 - j Differences in three properties – solubility, acidity and volatility – are used to remove impurities from the 2-chloro-2-methylpropane. Group the impurities removed under these headings.
 - k Work out the **% yield** of 2-chloro-2-methylpropane as follows:
 - i Write the balanced equation for the reaction used to produce 2-chloro-2-methylpropane .
 - ii What is the maximum mass of 2-chloro-2-methylpropane that you could obtain from the mass of 2-methylpropan-2-ol used?
 - iii Calculate the % yield you obtained in your experiment.
 - l Experience suggests that you will do well to get a yield over 50%. What reasons can you think of to account for a yield well under 100%?
 - m Look carefully at the equation for the preparation of 2-chloro-2-methylpropane. What type of reaction is this?
 - n What problems would you expect at each stage of the process if you were trying to scale it up to make several *tonnes* of 2-chloro-2-methylpropane?

In this activity you will analyse data about a number of compounds which have been considered as possible refrigerants. You will then investigate various replacements for CFCs and look at the international agreements that govern the phasing out of ozone-depleting compounds.

In the final part you are asked to produce a briefing document for a politician, which will allow you to practise your research and communication skills.

Part 1: Properties of refrigerants

Figure 1 shows an outline of how a refrigerator works.

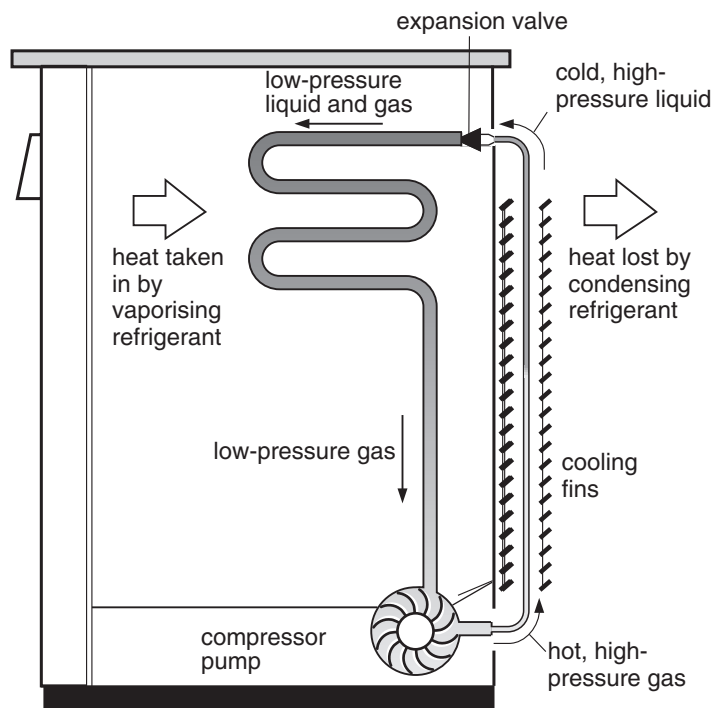


Figure 1 How a refrigerator works

The refrigerant fluid is crucial to the whole process. It must have the right boiling point – high enough to liquefy by compression, low enough to vaporise easily at reduced pressure; 240 K (–33 °C) is about right. For many years, ammonia was used as a refrigerant.

a What other properties do you think a refrigerant fluid should have?

Compounds derived from alkanes by replacing H atoms with Cl and F atoms appeared to be good candidates for refrigerant fluids. The different halogen atoms give the compounds different properties.

Table 1 gives some properties of alkanes, chloroalkanes and fluoroalkanes. Suggest answers to the following questions concerning the properties of the compounds in the table.

Formula	Boiling point/K	Chemical reactivity in atmosphere	Flammable?	Toxic?	Price
CH ₄	109	moderate	yes	no	low
CCl ₄	350	low	no	yes	medium
CF ₄	144	very low	no	no	high
C ₂ H ₆	185	moderate	yes	no	low
C ₂ Cl ₆	459	low	no	yes	medium
C ₂ F ₆	194	very low	no	no	very high
C ₃ H ₈	231	moderate	yes	no	low
C ₃ Cl ₈	559	low	no	yes	high
C ₃ F ₈	237	very low	no	no	very high

Table 1 Properties of some alkanes, chloroalkanes and fluoroalkanes

- b Imagine you are in the position of Thomas Midgley in 1930 and are looking for a suitable replacement for ammonia as a refrigerant fluid. You are investigating the properties of hydrocarbons and halogenoalkanes.

What are the advantages and disadvantages of each of the following in molecules of a refrigerant fluid:

- H atoms?
- Cl atoms?
- F atoms?

Why do you think Midgley selected CCl_2F_2 as a suitable replacement for ammonia?

Part 2: Replacing CFCs

It may help to work in small groups for this part of the activity and then compare your findings with other groups in your class.

In the late 1980s, it was shown convincingly by scientists that CFCs being used for a wide variety of purposes were damaging the ozone layer in the stratosphere. It was vital therefore that replacements were found. Just prior to the acceptance of this evidence, the annual world production of CFCs was about 8.5×10^5 tonnes.

The tables below give information about compounds containing C, H, Cl and F atoms. The three CFCs that were most widely used are shown in Table 2. Table 3 gives information about possible replacement compounds. A key to the symbols and terms used in the tables is given below the tables.

Table 2 CFCs that were widely used

Formula	Code	Uses	Toxicity	Boiling point/K	Flammable?	Ozone depletion potential	Price
CCl_3F	CFC 11	PRB*	low	297	no	1.0	medium
CCl_2F_2	CFC 12	PRB*	low	243	no	1.0	medium
$\text{CCl}_2\text{FCClF}_2$	CFC 113	C*	low	321	no	0.8	high

* see key below

Table 3 Possible replacement compounds

Formula	Code	Toxicity	Boiling point/K	Flammable?	Ozone depletion potential	Price
CH_2Cl_2		high	313	no	$<<0.01$	low
$\text{CH}_3\text{CH}_2\text{CH}_3$		low	231	yes	0	low
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_3$		low	273	yes	0	low
CH_3OCH_3		low	249	yes (very)	0	low
CHClF_2	HCFC 22	low	232	no	0.06	high
$\text{CF}_3\text{CH}_2\text{F}$	HFC 134a	low	247	no	0	very high
$\text{CF}_3\text{CCl}_2\text{H}$	HCFC 123	low	302	no	0.02	high
$\text{CH}_3\text{CCl}_2\text{F}$	HCFC 141b	low	305	yes	0.11	high
CH_3CHF_2	HFC 152a	?low	249	yes	0	high

Key to symbols used in the tables:

- The code is the number used in industry to identify the compound.
HCFC: hydrochlorofluorocarbons
HFC: hydrofluorocarbons.
- R = refrigeration and air conditioning; B = foam blowing agent; P = aerosol propellant (a minor use); C = cleaning solvent.
- The ozone depletion potential (ODP) is a measure of the effectiveness of the compound in destroying stratospheric ozone. CCl_3F is defined as having an ODP of 1.0.

QUESTIONS

- c Hydrogen in the molecule results in it being degraded in the troposphere. Fluorine results in stability or inertness.
- Look at the ozone depletion potentials (ODPs) of the various compounds.
- What features of a compound seem to give rise to a high ODP? Suggest a reason why.
 - What features give rise to a low ODP? Suggest a reason why.
- Compare your answers and ideas with those of the other groups.
- d On the basis of the information in the tables, which compound do you consider to be the most suitable replacement for:
- CCl_3F (CFC 11), for use as a blowing agent?
 - CCl_2F_2 (CFC 12), for use as a refrigerant?
 - $\text{CCl}_2\text{FCClF}_2$ (CFC 113), for use as a cleaning solvent?

In each case, make a shortlist of two or three, then make your final selection. Give reasons for your choices.

What further information would you need before making a final decision about replacement compounds?

Part 3: International agreements

The first major international agreement which limited production of CFCs and other ozone-depleting compounds was the *Montreal Protocol*. It was signed by thirty countries in 1987 and came into force on 1 January 1989.

The Montreal Protocol

Production of CFCs 11, 12, 113, 114, 115:

- frozen at 1986 level from 1992
- reduced to 80% from 1993
- reduced to 50% from 1998.

Production of halons (brominated halogenoalkanes used in fire extinguishers):

- restricted to 1986 levels

In 1989 the European Council of Ministers announced that CFC production would be phased out completely by the year 2000. The US made a similar declaration. Later in the year the Montreal Protocol countries (then increased to more than 60) adopted the same policy.

The Revised Montreal Protocol, London Amendment, 1990

Production of CFCs:

- reduced to 50% from 1995
- reduced to 15% from 1997
- zero by 2000.

Tetrachloromethane (used as a solvent):

- reduced to 15% from 1995
- zero by 2000.

Production of halons:

- phased out by 2000 (with some exceptions as there are no known alternatives).

1,1,1-trichloroethane (used as a solvent):

- reduced to 70% from 1995
- reduced to 30% from 1997
- zero by 2005.

HCFCs, with ODPs of about 0.05, should be used to replace CFCs in about 15% of applications, but this is seen as a transitional measure with an HCFC phase-out expected for about 2030.

1 The Montreal protocol is still under review and continues to be amended.

Use Table 4 and the Internet to find out about the latest agreements for CFCs, halogenoalkanes and HCFCs. To start the search, use the Salters Advanced Chemistry web site.

2 Prepare a briefing document for a politician which:

- explains the relative merits, as refrigerants, of CFCs, HCFCs, alkanes and HFCs in terms of their constituent elements
- provides an update of international agreements on the use of CFCs and HCFCs
- identifies any outstanding problems you are aware of concerning the impact of international agreements.

Table 4 Reductions in consumption of some halogen-containing compounds required by the regulations

Provisions of the Montreal Protocol agreed and in force at the end of 1997					European Union regulations enacted and in force			
	CFCs	'Other' CFCs	Halons	HCFCs	CFCs	'Other' CFCs	Halons	HCFCs
Base Year	1986	1989	1986	1989	1986	1986	1986	1989
Year								
1993		20%			50%	50%	freeze	
1994	75%	75%	100%		85%	85%	100%	
1995					100%	100%		
1996	100%	100%		freeze				
2004				35%				35%
2007								60%
2010				65%				80%
2013								95%
2015				90%				100%
2020				99.5%				
2030				100%				

(Based on data given, by kind permission, in R. Powell, *Methods of Organic Chemistry* Vol. E10a, Organo-fluorine compounds, eds. B. Baasner, H. Hagermann, J.C. Tetlow. G. Thienne Verlag, 1999.)

This activity will help you record some of the information that has been presented in sections A1–A5 of this unit.

Now is the time to make sure you have a record of the main points covered in sections **A1–A5**. It is important to make these notes *as you go along*.

The following activities should help you check that you have understood the important ideas.

- a Make sure you know the approximate composition of the atmosphere and the main pollutant gases.
- b You have come across a lot of new terms in this part of the course, eg *stratosphere*, *radical*, *photodissociation*, *termination*. Go through the text and pick out these and others like them, then write a few lines of explanation for each, including an example taken from the text. (Don't forget to look in the **Activities** as well as in the **Storyline** and the **Chemical Ideas**.)
- c Make a summary of the key points in the story concerning the function of ozone in the stratosphere, and why there is concern about the effect CFCs have on this ozone. Your summary should be no more than half a page long.
- d Summarise the chemical reactions taking place in the stratosphere. You may find it helps to do this in the form of a diagram.

Which are the greenhouse gases?

In this activity, you will use information about the infrared absorption of atmospheric gases to decide which will contribute to the greenhouse effect.

The frequency and wavelength ranges over which certain important atmospheric gases absorb infrared radiation are shown in Table 1.

Gas	Absorption frequency range / 10^{14} Hz	Absorption wavelength range /nm
N ₂	none	none
O ₂	none	none
H ₂ O	0.43–0.64	6950–4670
CO ₂	0.18–0.24	17 000–12 500
CH ₄	0.39–0.46 and 0.85–1.03	7700–6500 and 3520–2900
N ₂ O	0.38–0.42	7800–7120
CCl ₃ F	0.25–0.33	12 000–9060
O ₃	0.28–0.32	10 700–9340

Notes: N₂O, dinitrogen oxide, is released by many organisms. Increased use of fertilisers increases the release of N₂O.

Table 1 Infrared absorption of some atmospheric gases

The spectrum in Figure 1 shows the infrared radiation emitted by the Earth. It shows the energy emitted at each wavelength (and frequency). Use the information in the table to answer the following questions.

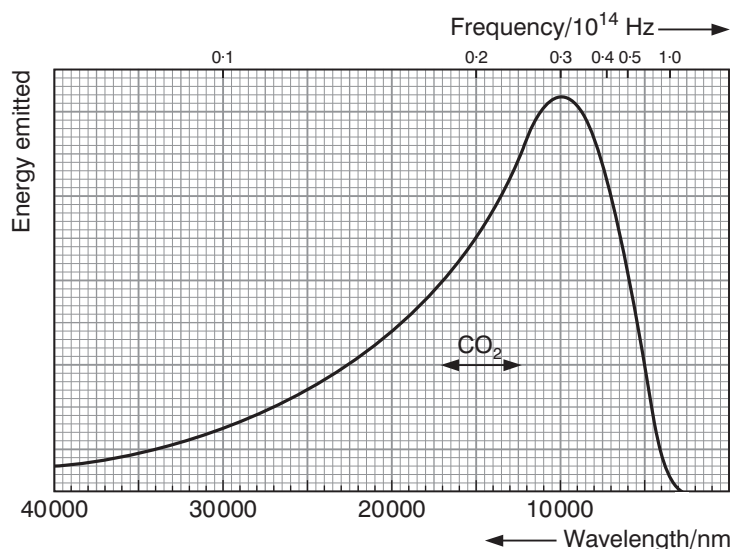


Figure 1 Infrared radiation emitted by the Earth

- The range of wavelengths absorbed by CO₂ in the atmosphere is shown in Figure 1. Mark on the Earth's radiation spectrum the absorption ranges (if any) of other gases that occur naturally in the atmosphere in considerable quantities: N₂, O₂, H₂O. (It is easier to use the wavelength ranges rather than the frequency ranges.)
- Which of these four *naturally abundant* gases are likely to have a greenhouse effect?
- What range of wavelengths of infrared radiation will be able to escape from Earth without absorption by these natural greenhouse gases? This range of frequencies acts as a 'window' which keeps the Earth cool by allowing radiation to escape.
- Now look at the other gases in the table: CH₄, N₂O, CCl₃F and O₃.
 - Which are produced naturally?
 - Which are produced as a result of human activities?
- Mark on the spectrum the absorption ranges of each of the gases mentioned in d.
 - Which are greenhouse gases?
 - Which absorb radiation in the 'window' mentioned in c?
- From the information given, it is not possible to compare the *magnitude* of the different gases' greenhouse effects. What further information would you need in order to do this?

A8.1

The effect of concentration changes on chemical equilibria

In this activity you will carry out some quick test-tube reactions to investigate the effect of changing the concentration of a substance in an equilibrium mixture.

Requirements

- test-tubes and rack
- potassium chromate(VI) solution, 0.1 mol dm^{-3} (2 cm^3)
- protective gloves
- dilute sulphuric acid, 1.0 mol dm^{-3} (2 cm^3)
- dilute sodium hydroxide solution, 2.0 mol dm^{-3} (2 cm^3)
- teat pipettes
- distilled water
- potassium (or ammonium) thiocyanate solution KSCN (or NH_4SCN), 0.5 mol dm^{-3} (2 drops)
- iron(III) chloride solution, 0.5 mol dm^{-3} (2 drops)
- solid ammonium chloride (1 spatula load)
- glass stirring rod

CARE Chromates(VI) irritate the skin and are suspected carcinogens. Avoid all skin contact. Any spillage should be washed off at once. Wear protective gloves.

CARE Thiocyanates and ammonium chloride are harmful.

ammonium chloride



HARMFUL

dilute sulphuric acid



IRRITANT

potassium chromate(VI) solution



TOXIC

potassium thiocyanate solution



HARMFUL

sodium hydroxide solution



CORROSIVE

CARE Eye protection and gloves must be worn.



WEAR EYE PROTECTION



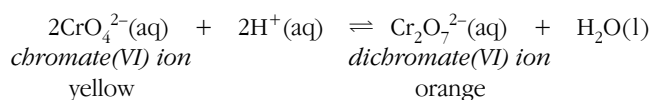
WEAR PROTECTIVE GLOVES

Introduction

The two equilibria in this activity involve coloured ions, so you can use the colour of the solution to investigate how the position of equilibrium changes when the concentration of one of the reactants is increased or decreased.

Equilibrium 1

Potassium chromate(VI) solution contains yellow chromate(VI) ions. In acidic solution, orange dichromate(VI) ions are formed.



- 1 Place 2 cm^3 potassium chromate(VI) solution (0.1 mol dm^{-3}) in a test-tube.
(**CARE** Chromates(VI) irritate the skin and are suspected carcinogens. Avoid all skin contact. Wear gloves.)
- 2 Add dilute sulphuric acid (**CARE** Irritant) drop by drop, with gentle shaking, until there is no further colour change.
- 3 Now add dilute sodium hydroxide solution (**CARE** Corrosive) drop by drop, with gentle shaking, until there is no further colour change. (Hydroxide ions, $\text{OH}^-(\text{aq})$, remove $\text{H}^+(\text{aq})$ ions from the solution.)
- 4 Repeat steps 2 and 3, and then record your observations in Table 1. Suggest a cause for each colour change (in terms of the concentration of the coloured ions) and then say what this tells you about the change in the position of equilibrium.

Change	Observation	Cause	What has happened to the position of equilibrium?
concentration $\text{H}^+(\text{aq})$ increased			
concentration $\text{H}^+(\text{aq})$ decreased			

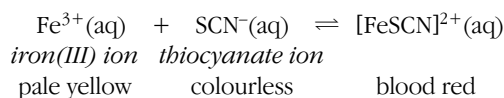
Table 1

QUESTION

- a Write an ionic equation to show how hydroxide ions, $\text{OH}^-(\text{aq})$, remove $\text{H}^+(\text{aq})$ ions from the solution. What type of reaction is this?

Equilibrium 2

Iron(III) ions and thiocyanate ions react in solution to produce a blood-red compound:



- Mix together 1 drop of iron(III) chloride solution (0.5 mol dm^{-3}) and 1 drop of potassium thiocyanate solution (0.5 mol dm^{-3}) in a test-tube, and add about 5 cm^3 of distilled water to form a pale orange-brown solution.
- Divide this solution into four equal parts in four test-tubes.
- Add 1 drop of iron(III) chloride solution to one test-tube, and 1 drop of potassium thiocyanate solution to a second. Compare the colours of these solutions with the two remaining tubes. Enter your observations in Table 2.
- Add a spatula-load of solid ammonium chloride to a third test-tube and stir well. (The effect of this is to reduce the concentration of iron(III) ions in the solution. The chloride ions in ammonium chloride react with the $\text{Fe}^{3+}(\text{aq})$ ions to produce $[\text{FeCl}_4]^-$ ions.) Compare the colour of this solution with the remaining tube and note your observation. Now complete the rest of Table 2.

Change	Observation	Cause	What has happened to the position of equilibrium?
concentration $\text{Fe}^{3+}(\text{aq})$ increased			
concentration $\text{SCN}^-(\text{aq})$ increased			
concentration $\text{Fe}^{3+}(\text{aq})$ decreased			

Table 2

QUESTION

- b How would the position of equilibrium be affected by increasing the concentration of $[\text{FeSCN}]^{2+}(\text{aq})$ ions?

SUMMARY

- c How does changing the concentration of one of the substances in an equilibrium mixture affect the position of equilibrium? Write a short paragraph to summarise your findings.

Measuring the concentration of carbon dioxide in air samples

In this activity you will see how pH measurements can be used to find the concentration of carbon dioxide in samples of air. You will then consider the limitations of the method.

Introduction

To understand this activity you need to know what is meant by the pH of a solution. You will learn about this in more detail in **The Oceans**. For now it is sufficient to know the following.

The pH scale is a measure of the acidity or alkalinity of a solution. Pure water has $\text{pH} = 7$. Solutions with $\text{pH} < 7$ are acidic; solutions with $\text{pH} > 7$ are alkaline.

The pH scale represents the concentration of $\text{H}^+(\text{aq})$ ions in a solution on a *logarithmic scale*. This means that a *tenfold* change in the concentration of $\text{H}^+(\text{aq})$ ions is needed to produce a pH change of 1.

This is shown in Figure 1, which gives the hydrogen ion concentrations, $[\text{H}^+(\text{aq})]$, corresponding to some pH numbers.

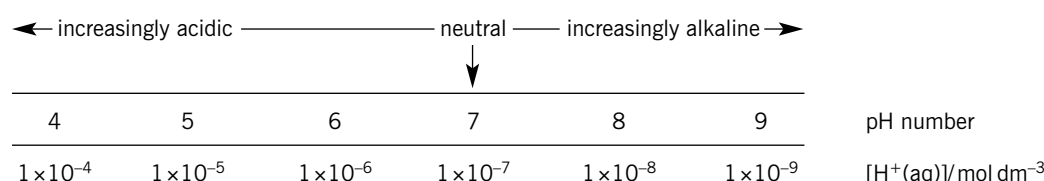


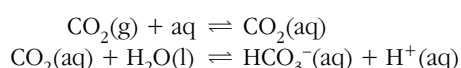
Figure 1 The pH scale

If the concentration of $\text{H}^+(\text{aq})$ ions in a solution is halved, the pH only changes by 0.3.

Using pH to measure CO_2 concentration in air

Carbon dioxide is a weakly acidic gas, and this property can be used to measure the concentration of carbon dioxide in samples of air.

The carbon dioxide in an air sample is allowed to dissolve in water. Then the acidity of the water is measured. The greater the carbon dioxide concentration in the air sample, the more acidic the water will be. Here are the important reactions:



However, changes in the concentration of dissolved carbon dioxide lead to only small changes in the pH of the solution.

If a solution of sodium hydrogencarbonate is used in place of water, the pH of the solution also changes with the concentration of carbon dioxide in the air sample, but the changes in pH are larger, so that the method becomes much more sensitive.

A pump is used to bubble air through a solution of sodium hydrogencarbonate until equilibrium is reached (about 20 minutes). The air can be drawn from different sources for comparison (eg outdoors and indoors). Some sample results are shown in Table 1.

The calibration graph (Figure 2) allows you to relate the pH of the sodium hydrogencarbonate solution to the concentration of carbon dioxide in the air sample.

(Biologists frequently use this technique to measure carbon dioxide concentrations in air samples. Instead of using a pH meter, they often use *bicarbonate indicator*. This is a standard solution of sodium hydrogencarbonate containing a mixture of indicators. These change colour according to the pH of the solution.)

Sample	pH
Outdoor air	9.3
Indoor air	8.7
Carbon dioxide	5.9

Table 1 Sample results (bubbling the gas into the solution for 20 minutes)

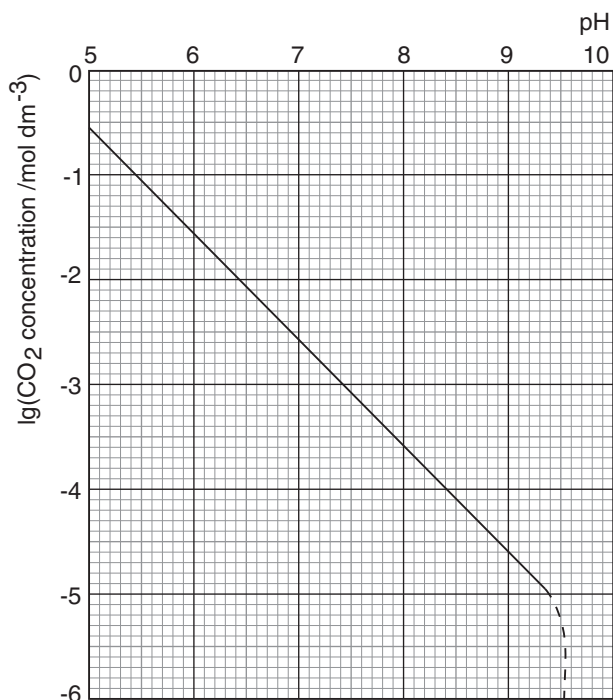


Figure 2 Calibration graph relating pH of the sodium hydrogencarbonate solution to the logarithm of the concentration of carbon dioxide in the air sample

QUESTIONS

- Use the sample results and the calibration graph to calculate the concentration of gaseous carbon dioxide in each sample. (Assume that equilibrium has been reached between the sample and the solution.)
- What is the uncertainty associated with a pH value which is read to 0.1 pH units? Calculate the percentage error for each of the pH values in the sample results.
- A difference of 0.1 pH units produces quite large differences after using the calibration graph and calculating the concentration
eg pH 9.0 gives a concentration of $2.5 \times 10^{-5} \text{ mol dm}^{-3}$
but pH 8.9 gives a concentration of $3.6 \times 10^{-5} \text{ mol dm}^{-3}$
What would be the percentage difference in the concentration calculated if the correct reading was pH 9.0 but the recorded value was 8.9?
- Is this method sensitive and accurate enough to detect the difference in carbon dioxide concentration between:
 - normal air and carbon dioxide gas?
 - air outside and inside the laboratory?
- Comment on the limitations of this method as a way of finding the concentration of carbon dioxide in the atmosphere. How useful would the method be for chemists monitoring pollution of the atmosphere?
- Your results for the concentration of carbon dioxide in the air samples are in mol dm^{-3} . The more usual measure of gaseous concentration is percentage by volume. Convert your results into percentage by volume. (Assume that at room temperature and pressure 1 mole of carbon dioxide occupies 24 dm^3 .)

Reducing the rate of increase of carbon dioxide in the atmosphere is one of the greatest challenges for the world as we move into the 21st century. In this activity you will write a report using information from a variety of sources and try your hand at writing an abstract which summarises the content of your report.

Introduction

Carbon dioxide is the greenhouse gas causing most concern. Increased carbon dioxide emissions could upset the existing balance in the atmosphere.

Below, four possible approaches to the control and reduction of atmospheric carbon dioxide are outlined:

Approach 1 Choose the right fuel.

Approach 2 Economise on the use of energy in general.

Approach 3 Dispose of the carbon dioxide.

Approach 4 Encourage photosynthesis.

Your teacher will probably tell you which of these approaches to work on, but first answer the questions below.

Table 1 gives details of four possible fuels.

Fuel	Typical constituent	Standard enthalpy change of combustion of typical constituent/kJ mol ⁻¹
coal	carbon	-394
natural gas	methane, CH ₄	-890
gasoline	octane, C ₈ H ₁₈	-5470
hydrogen	hydrogen, H ₂	-286

Table 1 Four possible fuels

- For each fuel in Table 1, write an equation to represent its combustion.
- For each fuel, calculate the amount in moles of carbon dioxide given out per kJ of energy released by the fuel.
- What are the implications of your answers?

What you do

- Write a report of no more than 400 words on *one* of the four approaches.

Refer to the statistics on the **Information Sheet** (*World resources and the consumption of energy*) as appropriate, and use any other sources of information you can find. (These may be Salters Advanced Chemistry units, textbooks, articles or the Internet.) At the end of your report, list clearly any additional sources you have used. (This list is not included in the word count.) You should refer to this list in your report where appropriate. For example, where you have copied material directly from another source, this must be indicated in the text and the source properly acknowledged.

You may be asked to make an oral presentation to the class on your approach.

- In the second part of this activity, a person is chosen to present a short briefing to the class on each of the approaches. After all four briefings have been presented, the class can work together to discuss an overall approach – which might be a combination of some or all of the four.

The overall objective is to recommend a strategy for bringing about a substantial reduction in the amount of carbon dioxide emitted into the atmosphere throughout the world.

The scale of the problem

When tackling this activity, you must be aware of the scale of the problem. It is enormous, because the amount of fossil fuel burned throughout the world is now huge: about 30 thousand million tonnes (30 Gt) of carbon dioxide are released into the atmosphere each year from human activities. The problem is much too big for a quick, easy solution.

You should also be aware of the scale of natural processes involving carbon dioxide compared with the scale of carbon dioxide production by humans. To remind yourself, have a look at the diagram of the global carbon cycle in

Storyline A8.

Approach 1: Choose the right fuel

Some fossil fuels give out more carbon dioxide than others when they burn.

Prepare a report on the control of carbon dioxide by selective use of fuels. You should decide which fuels are preferable for which uses. You should indicate any problems and drawbacks involved in this approach.

Bear in mind the following points:

- Your scheme will have to take account of the available reserves of the fuels concerned.
- Different fuels are appropriate for different uses. For example, most forms of transport need liquid fuels.
- Electricity generation is a major use of fossil fuels.
- Hydrogen is a possible fuel that may have attractions, but hydrogen does not occur naturally: it has to be manufactured.

You might also consider the implications of using fuels derived from biomass (eg ethanol produced from grain or from wood) rather than those from fossil fuels.

Approach 2: Economise on the use of energy in general

All fossil fuels release carbon dioxide when they burn. Economising on the use of energy in general will reduce consumption of fossil fuels, and so reduce carbon dioxide emission. Energy consumption can also be reduced by increasing the efficiency of energy conversions, particularly in power stations and motor vehicle engines.

Prepare a report on the control of carbon dioxide by economising on the use of energy worldwide. You should make suggestions concerning how this could be done. You should indicate any problems and drawbacks involved in this approach. Try to find information about some of the techniques for increasing fuel efficiency, for example fluidised bed combustion in power stations.

Bear in mind the following points:

- Standard of living is closely related to energy consumption (but see the next point).
- There are examples, over the last 30–40 years, of energy economy and efficiency drives in Western countries. These have led to a reduction in the use of fossil fuels without causing a significant fall in standard of living.
- Electricity generation is a major use of fossil fuels.
- You may want to consider alternative energy sources. Nuclear energy may have attractions in this context, but bear in mind that it carries with it a number of environmental hazards.

Approach 3: Dispose of the carbon dioxide

It might be possible to get rid of carbon dioxide, after it has been formed but before it can escape into the atmosphere. The problem with this approach is the enormous quantity of carbon dioxide involved: 30 thousand million tonnes each year.

Carbon dioxide removal is not yet a practical possibility on a large scale, though a number of methods have been suggested. One possibility is to use a special solvent in which carbon dioxide is highly soluble. Exhaust gases from a power station would be passed through the solvent so that the carbon dioxide dissolved. The solvent would then be heated to expel the carbon dioxide. The solvent would be recycled.

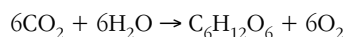
The carbon dioxide expelled from the solvent would then have to be disposed of. The question is, where? It could in theory be liquefied and pumped to a depth of more than 500 metres beneath the sea. At this depth the pressure is so great that the carbon dioxide would remain at the bottom as a liquid. Another possibility would be to dump the carbon dioxide underground in spent oil or gas wells from which all the oil or gas had been removed.

The technology for this kind of process is expensive. It might double the capital cost of a small coal-fired power station, and it could raise the price of electricity by about 75%.

Prepare a report on the control of carbon dioxide by disposing of the gas. You should indicate the advantages of this approach and mention any problems and drawbacks involved. If you can invent other possible ways of absorbing and disposing of carbon dioxide, you could mention these too.

Approach 4: Encourage photosynthesis

Photosynthesis removes carbon dioxide from the atmosphere, converting it to carbohydrate and oxygen.



Worldwide, about 500 Gt of carbon dioxide are removed by photosynthesis each year: compare this with the 30 Gt produced from the combustion of fossil fuels.

Increasing the rate of global photosynthesis will help remove carbon dioxide from the atmosphere. Photosynthesis occurs particularly rapidly in the rain forests of the tropics. However, the rain forests are being destroyed at a great rate by human activities. What is more, destroying the forests actually adds more carbon dioxide to the atmosphere when the wood is burned or decays. It is estimated that about 2 Gt of carbon dioxide per year are released due to the destruction of forests around the world.

At present, photosynthesis can only be carried out by plants. But a great deal of research is going on to try and find out the exact mechanism by which photosynthesis works. Once this is understood, it may be possible to arrange for photosynthesis to occur in artificial systems, outside plants. There might then be the possibility of using artificial photosynthesis to absorb atmospheric carbon dioxide and at the same time produce carbohydrates for food or fuels.

Prepare a report on the control of carbon dioxide by encouraging photosynthesis. You should indicate the advantages of this approach and mention any problems and drawbacks involved. If you can think of other possible ways of making use of photosynthesis, you could mention these too.

Information Sheet: World resources and the consumption of energy

Energy source	Consumption /million tonnes of oil equivalent*
oil	3389
natural gas	2016
coal	2219
hydroelectric	226
nuclear	627
total	8477

Table 1 World consumption of different energy sources, 1998 (BP Amoco Statistical Review of World Energy, 1999)

Fuel	Reserves /million tonnes of oil equivalent*
oil	143 000
natural gas	134 000
coal	656 000
total	963 000

Table 2 Estimated world reserves of fossil fuels, 1998 (BP Amoco Statistical Review of World Energy, 1999)

Region	Consumption/million tonnes of oil equivalent*					
	Oil	Gas	Coal	Hydro	Nuclear	Total
North America	1008	647	566	57	204	2492
South and Central America	217	77	19	45	3	361
Europe	760	385	351	50	243	1788
Former Soviet Union	184	746	167	20	50	897
China	190	17	615	17	4	843
Japan	255	63	88	9	84	499
Africa	112	44	96	6	4	262

Table 3 Energy consumption in different regions of the world, 1998 (BP Amoco Statistical Review of World Energy, 1999)

Fuel	Percentage of generated electricity
coal	38.4
oil	9.3
gas	14.8
nuclear	17.7
hydroelectric	18.4
other (inc. geothermal, solar)	1.4

Table 4 Percentage of the electricity generated from different fuels worldwide in 1996 (Key World Energy Statistics, IEA, 1998)

Fuel	Percentage of fuel consumed
coal	24.0
oil	35.3
gas	20.2
nuclear	6.7
hydroelectric	2.3
combustible residues and waste	11.1
other (inc. geothermal, solar)	0.4

Table 5 World fuel consumption in 1996 (Key World Energy Statistics, IEA, 1998)

* So that there can be an easy comparison between fuels, the consumption is calculated in terms of the energy generated by oil: one tonne of oil equivalent equals 42 GJ or, for example, 1.5 tonnes of hard coal. It produces 12 MW hr of electricity in a modern power station.

Check your notes on The Atmosphere

This activity helps you get your notes in order at the end of this unit.

Use this list as the basis of a summary of the unit by collecting together the related points and arranging them in groups. Check that your notes cover the points and are organised in appropriate ways. Remember that you will be coming back to many of the ideas in later units.

Most of the points are covered in the **Chemical Ideas**, with supporting information in the **Storyline** or **Activities**. However, if the *main* source of information is the Storyline or an Activity, this is indicated.

- The gases present in the atmosphere, including some major pollutants; understand values for composition by volume measured in percentage concentration and in parts per million (ppm) (**Storyline A1**).
- The idea that rotational, vibrational and electronic energies are quantised.
- The qualitative changes in rotational, vibrational and electronic energy of molecules caused by the absorption of radiation of appropriate frequency.
- The relationship between frequency and energy of electromagnetic radiation.
- The structure and reactivity of ozone and the way it is formed and destroyed in the stratosphere; how ozone acts as a sunscreen (**Storyline A3**; **Activity A3.1**).
- The factors that affect the rate of a chemical reaction and the use of collision theory to explain the effects.
- The meaning of the terms: *enthalpy profile* and *activation enthalpy*.
- The use of the concept of activation enthalpy to explain the qualitative effect of temperature changes on rate of reaction.
- The role of catalysts in providing alternative routes of lower activation enthalpy; homogeneous catalysis in terms of the formation of intermediates.
- The difference between homolytic and heterolytic fission of a covalent bond.
- The formation, nature and reactivity of radicals; the mechanism of a radical chain-reaction involving initiation, propagation and termination.
- The reaction of alkanes with halogens (**Activities A3.2** and **A3.3**).
- The nature and names of halogenoalkanes.
- The meaning of the terms: *hydrolysis*, *substitution*, *nucleophile* and *carbocation*.
- Outline of the preparation of a halogenoalkane from an alcohol and the principle stages in the purification of an organic liquid product (**Activity A4.2**).
- The characteristic properties of halogenoalkanes, comparing fluoro-, chloro-, bromo- and iodo-compounds: boiling points, formation of radicals by interaction with ultraviolet radiation (**Storyline A3**), and nucleophilic substitution with water, hydroxide ions and ammonia.
- The mechanism of nucleophilic substitution in halogenoalkanes.
- The use of relative electronegativity values to predict bond polarity in a covalent bond; the relationship between reactivity of halogenoalkenes and bond enthalpy and bond polarity.
- The nature and uses of chlorofluorocarbons (CFCs) (**Storyline A4**) and the relative advantages and disadvantages of replacement compounds (**Activity A4.3**).
- The chemical basis of the depletion of ozone in the stratosphere due to halogenoalkanes, involving the formation of halogen atoms and the catalytic role of these atoms in ozone destruction (**Storyline A3**).
- The relationship between the 'greenhouse effect' in the troposphere and the absorption characteristics of atmospheric gases (**Storyline A6** and **A7**; **Activity A6**).
- A comparison of the different approaches to the control of global warming through the control of carbon dioxide emissions (**Storyline A9**; **Activity A9**).
- The meaning of the term: *dynamic equilibrium*.
- The physical and chemical changes occurring when carbon dioxide dissolves in water, and the associated equilibria.
- The use of Le Chatelier's Principle to explain and predict the effects on the position of equilibrium of changes in concentration, temperature and pressure.